

Classification of materials



For us, pure actually means having no adulteration.

But for a scientist pure means that it contains only one type of matter or particles all of which are same in their chemical nature.



Is Matter Around us Pure

- ☐ Introduction
- ☐ Type of matter
- ☐ Types of mixture
- ☐ Solution
- ☐ Methods of expressing concentration of solutions
- ☐ Solubility
- ☐ Suspension
- ☐ Colloidal solution
- ☐ Separation of mixtures

INTRODUCTION

If we observe some **soil** and some **sugar** placed on two different sheets of paper with a magnifying glass.

1. We find that soil contains clay particles, some grass particles and even some dead insects, etc. That is, soil contains particles of different kinds, thus it is called an **impure substance**.
2. Sugar which contains only one kind of particles, is called a **pure substance**.



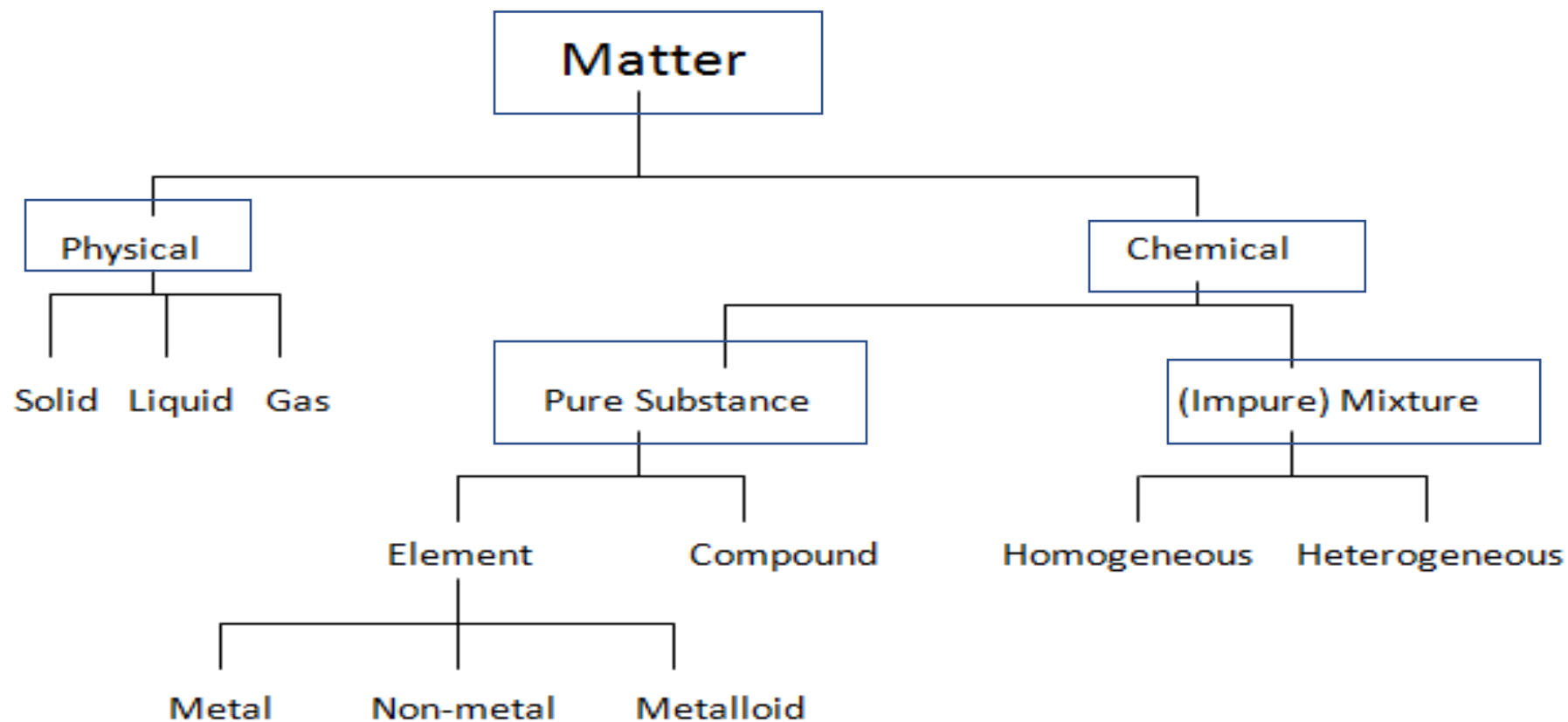
- A **substance** is a kind of matter which cannot be separated into any other types of matter by some physical means.
- Substance which has only one component and nothing else in it is called a **pure substance**.
- Substances are mostly mixed with one another and their combination is known as **mixture**.

What about milk????????????
Is it pure substance or a mixture????????

.



TYPE OF MATTER



Pure substance

A homogeneous substance which contains particles of only one kind and has definite set of properties is called a pure substance.

For example: Oxygen (O_2), Iron, sulphur, etc.

OR

If a substance is composed of two or more different kinds of particles combined together in fixed proportion by weight, then also substance is regarded as a pure substance.

For example: Sodium chloride($NaCl$), Magnesium oxide(MgO), Carbon dioxide(CO_2), Copper sulphate($CuSO_4$), etc.

Characteristics of pure substances

1. A pure substance is homogeneous in nature.
2. A pure substance has a definite set of properties.
3. The composition of a pure substance cannot be altered by any physical means.
4. A pure substance has a definite melting point, density, boiling point, etc.

For example:

Water is a pure substance because it has uniform composition throughout. It consists of only one kind of particles called water molecules, which cannot be separated into constituents(H and O) by any simple physical process.

Question

Can we consider all homogeneous substances as pure?

Explanation:

No, For example, common salt is solution in water, is a homogeneous solution. Yes it cannot be called as pure substance, as it is made of two different substances i.e. salt and water.



Mixture

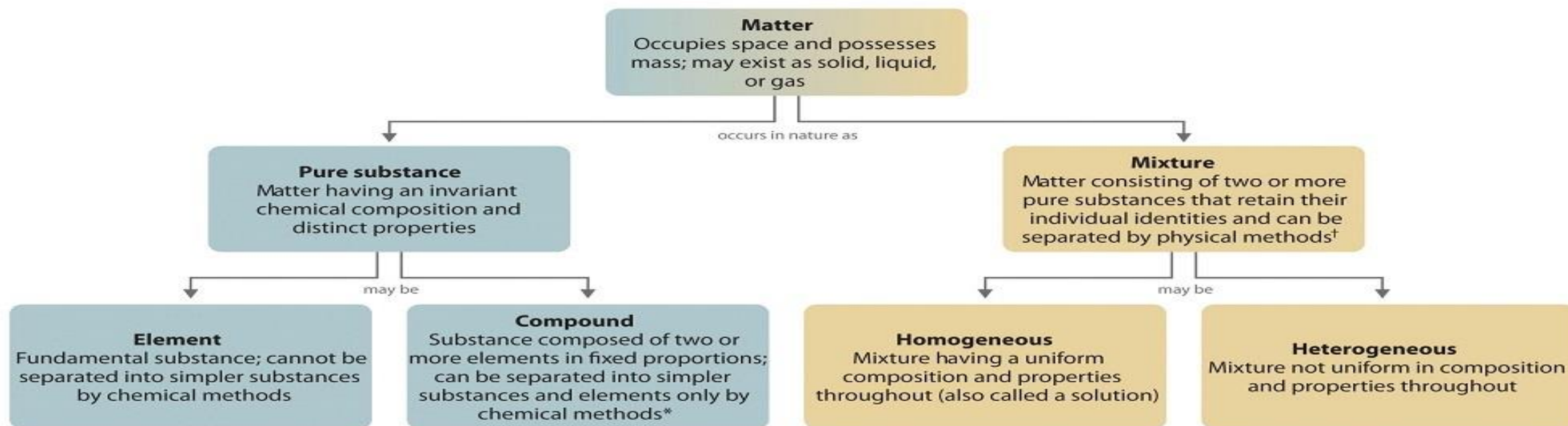
When two or more substances(elements, compounds or both) are physically mixed together in any proportion, such that they do not undergo any chemical change, but retain their individual characteristics, the resulting product is called a mixture.

For example:

1. Brass is mixture of copper and zinc.
2. Crude oil is a mixture of large number of different hydrocarbons.
3. If we dissolve sugar in water both of which are pure compounds, the solution of sugar in water is a mixture.

Characteristics of mixtures

1. A mixture may be homogeneous or heterogeneous.
2. The constituents of a mixture can be separated by physical means like filtration, evaporation, sublimation and magnetic separation.
3. In the preparation of a mixture energy is neither evolved nor absorbed.
4. A mixture has no definite melting point and boiling point.
5. The constituents of a mixture retain their original set of properties. For example, magnet attracts iron fillings in a mixture of sand and iron filings.



* Chemical methods of separation include electrolysis.

† Physical methods of separation include filtration, distillation, and crystallization.

Iron Metal
 $\text{Fe}_{(s)}$



Table Salt
 $\text{NaCl}_{(s)}$



Apple Juice



Salad Dressing



Question

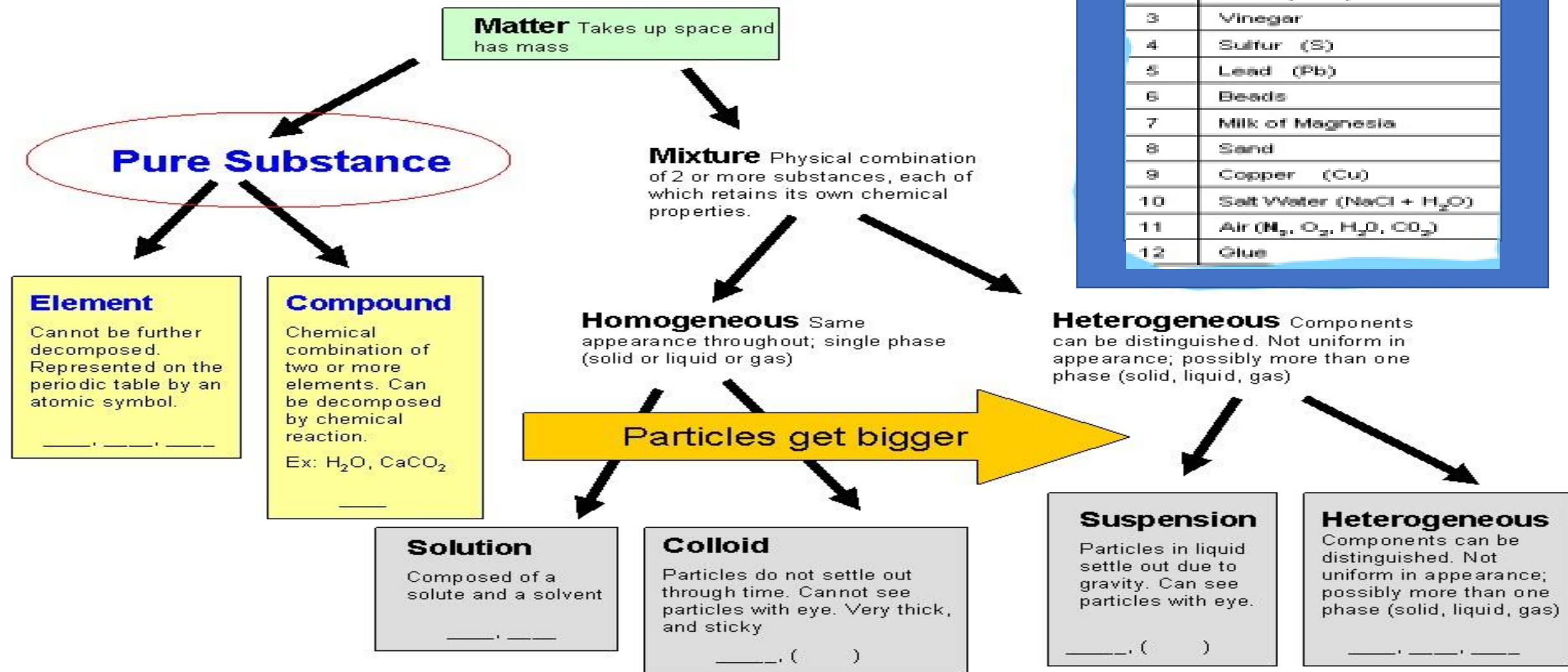
Why are composite materials often used instead of single material?

Explanation:

Composite material are those that combine the properties of two constituents. In order to get the exact properties needed for particular job. For example,

1. Steel: It is used to make rust free, unbreakable utensils.
2. Bone is composite material.
3. Wood consists of cellulose mixed with 'Lignin' which is largely responsible for the strength of the wood.

Directions: Classify each sample to its most specific form possible (furthest down on the paper). Write the number to the corresponding sample in the correct space provided.



Elements

Robert Boyle was the first scientist who used the term element in 1661. Antoine Laurent Lavoisier(1743-1794), a French chemist was the first to give a useful definition of an element as follows.

An element is defined as the simplest form or the basic form of a pure substance which can neither be broken into nor built up from simpler substances by any physical or chemical method. For example, iron is an element because we may break it into pieces, as small as possible, but the chemical formula of all the pieces will be same as that of iron.

Note

This definition given above is not considered to be absolutely correct now a days. This is because after the discovery of radioactivity, it has been found that elements can be broken into simpler substance. Furthermore, a number of elements have been synthesized from simpler substances. Hence, the definition of an element has been modifies as follows:

An element is now defined as a pure substance that contains only one kind of atoms.

An element is made up of only one kind of atoms which are identical in all respects, size, mass, composition, etc.

However, atoms of different elements differ in size, mass and composition.

118 Elements have been discovered so far. Out of these 90 elements have been found to occur in nature whereas the remaining 28 have been synthesized by the scientists in the laboratories by suitable chemical reactions and nuclear reactions.

Elements are found to exist in all the three states of matter:

Solid elements

Majority of elements are solids at room temperature. For example, iron, copper, aluminium, silver, gold, sodium, potassium, sulphur, phosphorus, carbon(graphite or diamond), iodine, etc.

Liquid elements

Mercury and bromine are the only two elements which are known to exist in the liquid state at room temperature. Gallium and caesium become liquid at a temperature slightly above room temperature(303K).

Gaseous elements

Eleven elements are gaseous at room temperature. These are hydrogen, oxygen, nitrogen, fluorine, chlorine, helium, neon, argon, krypton, xenon and radon.

Periodic Table of the Elements

Atomic Number → → Symbol Name → Electrons per shell → 1 H Hydrogen 1.008 1 ← Atomic Weight																																				18 VIIIA 2 He Helium 4.0026 2													
1 IA H Hydrogen 1.008 1		2 IIA Be Beryllium 9.0122 2-2																13 IIIA B Boron 10.81 2-3		14 IVA C Carbon 12.011 2-4		15 VA N Nitrogen 14.007 2-5		16 VIA O Oxygen 15.999 2-6		17 VIIA F Fluorine 18.998 2-7		10 Ne Neon 20.180 2-8																					
3 Li Lithium 6.94 2-1		4 Na Sodium 22.98976928 2-8-1		11 Li Lithium 6.94 2-1		12 Mg Magnesium 24.305 2-8-2		19 Na Sodium 22.98976928 2-8-1		20 Ca Calcium 40.078 2-8-8-2		37 Na Sodium 22.98976928 2-8-1		38 Rb Rubidium 85.4678 2-8-18-8-1		55 Cs Caesium 132.90545196 2-8-18-32-18-8-1		87 Fr Francium (223) 2-8-18-32-18-8-1		104 Rf Rutherfordium (267) 2-8-18-32-10-2		105 Db Dubnium (268) 2-8-18-32-11-2		106 Sg Seaborgium (269) 2-8-18-32-12-2		107 Bh Bohrium (270) 2-8-18-32-13-2		108 Hs Hassium (277) 2-8-18-32-14-2		109 Mt Meitnerium (278) 2-8-18-32-15-2		110 Ds Darmstadtium (281) 2-8-18-32-17-1		111 Rg Roentgenium (282) 2-8-18-32-17-2		112 Cn Copernicium (285) 2-8-18-32-18-2		113 Nh Nihonium (286) 2-8-18-32-18-3		114 Fl Flerovium (289) 2-8-18-32-18-4		115 Mc Moscovium (290) 2-8-18-32-18-5		116 Lv Livermorium (293) 2-8-18-32-18-6		117 Ts Tennessee (294) 2-8-18-32-18-7		118 Og Oganesson (294) 2-8-18-32-18-8	
State of matter (color of name) GAS LIQUID SOLID UNKNOWN																		Subcategory in the metal-metalloid-nonmetal trend (color of background) Alkali metals Alkaline earth metals Transition metals Lanthanides Actinides Post-transition metals Metalloids Reactive nonmetals Noble gases Unknown chemical properties																															
21 IIIB Sc Scandium 44.955908 2-8-9-2		22 IVB Ti Titanium 47.867 2-8-10-2		23 VB V Vanadium 50.9415 2-8-11-2		24 VIB Cr Chromium 51.9961 2-8-13-1		25 VIIB Mn Manganese 54.938044 2-8-13-2		26 VIIIB Fe Iron 55.845 2-8-14-2		27 VIIIB Co Cobalt 58.933 2-8-15-2		28 VIIIB Ni Nickel 58.693 2-8-16-2		29 IB Cu Copper 63.546 2-8-18-1		30 IIB Zn Zinc 65.38 2-8-18-2		31 Ga Gallium 69.723 2-8-18-3		32 Ge Germanium 72.630 2-8-18-4		33 As Arsenic 74.922 2-8-18-5		34 Se Selenium 78.971 2-8-18-6		35 Br Bromine 79.904 2-8-18-7		36 Kr Krypton 83.798 2-8-18-8																			
39 Y Yttrium 88.90584 2-8-18-9-2		40 Zr Zirconium 91.224 2-8-18-10-2		41 Nb Niobium 92.90637 2-8-18-12-1		42 Mo Molybdenum 95.95 2-8-18-13-1		43 Tc Technetium (98) 2-8-18-13-2		44 Ru Ruthenium 101.07 2-8-18-15-1		45 Rh Rhodium 102.91 2-8-18-16-1		46 Pd Palladium 106.42 2-8-18-18		47 Ag Silver 107.87 2-8-18-18-1		48 Cd Cadmium 112.41 2-8-18-18-2		49 In Indium 114.82 2-8-18-18-3		50 Sn Tin 118.71 2-8-18-18-4		51 Sb Antimony 121.76 2-8-18-18-5		52 Te Tellurium 127.60 2-8-18-18-6		53 I Iodine 126.90 2-8-18-18-7		54 Xe Xenon 131.29 2-8-18-18-8																			
57-71 Lanthanides		72 Hf Hafnium 178.49 2-8-18-32-10-2		73 Ta Tantalum 180.94788 2-8-18-32-11-2		74 W Tungsten 183.84 2-8-18-32-12-2		75 Re Rhenium 186.21 2-8-18-32-13-2		76 Os Osmium 190.23 2-8-18-32-14-2		77 Ir Iridium 192.22 2-8-18-32-15-2		78 Pt Platinum 195.08 2-8-18-32-17-1		79 Au Gold 196.97 2-8-18-32-18-1		80 Hg Mercury 200.59 2-8-18-32-18-2		81 Tl Thallium 204.38 2-8-18-32-18-3		82 Pb Lead 207.2 2-8-18-32-18-4		83 Bi Bismuth 208.98 2-8-18-32-18-5		84 Po Polonium (209) 2-8-18-32-18-6		85 At Astatine (210) 2-8-18-32-18-7		86 Rn Radon (222) 2-8-18-32-18-8																			
89-103 Actinides		104 Rf Rutherfordium (267) 2-8-18-32-10-2		105 Db Dubnium (268) 2-8-18-32-11-2		106 Sg Seaborgium (269) 2-8-18-32-12-2		107 Bh Bohrium (270) 2-8-18-32-13-2		108 Hs Hassium (277) 2-8-18-32-14-2		109 Mt Meitnerium (278) 2-8-18-32-15-2		110 Ds Darmstadtium (281) 2-8-18-32-17-1		111 Rg Roentgenium (282) 2-8-18-32-17-2		112 Cn Copernicium (285) 2-8-18-32-18-2		113 Nh Nihonium (286) 2-8-18-32-18-3		114 Fl Flerovium (289) 2-8-18-32-18-4		115 Mc Moscovium (290) 2-8-18-32-18-5		116 Lv Livermorium (293) 2-8-18-32-18-6		117 Ts Tennessee (294) 2-8-18-32-18-7		118 Og Oganesson (294) 2-8-18-32-18-8																			
57 La Lanthanum 138.91 2-8-18-19-9-2		58 Ce Cerium 140.12 2-8-18-19-9-2		59 Pr Praseodymium 140.91 2-8-18-21-8-2		60 Nd Neodymium 144.24 2-8-18-22-8-2		61 Pm Promethium (145) 2-8-18-23-8-2		62 Sm Samarium 150.36 2-8-18-24-8-2		63 Eu Europium 151.96 2-8-18-25-8-2		64 Gd Gadolinium 157.25 2-8-18-25-9-2		65 Tb Terbium 158.93 2-8-18-27-8-2		66 Dy Dysprosium 162.50 2-8-18-28-8-2		67 Ho Holmium 164.93 2-8-18-29-8-2		68 Er Erbium 167.26 2-8-18-30-8-2		69 Tm Thulium 168.93 2-8-18-31-8-2		70 Yb Ytterbium 173.05 2-8-18-32-8-2		71 Lu Lutetium 174.97 2-8-18-32-9-2																					
89 Ac Actinium (227) 2-8-18-32-18-9-2		90 Th Thorium 232.04 2-8-18-32-18-10-2		91 Pa Protactinium 231.04 2-8-18-32-20-9-2		92 U Uranium 238.03 2-8-18-32-21-9-2		93 Np Neptunium (237) 2-8-18-32-22-9-2		94 Pu Plutonium (244) 2-8-18-32-24-8-2		95 Am Americium (243) 2-8-18-32-25-8-2		96 Cm Curium (247) 2-8-18-32-25-9-2		97 Bk Berkelium (247) 2-8-18-32-27-8-2		98 Cf Californium (251) 2-8-18-32-28-8-2		99 Es Einsteinium (252) 2-8-18-32-29-8-2		100 Fm Fermium (257) 2-8-18-32-30-8-2		101 Md Mendelevium (258) 2-8-18-32-31-8-2		102 No Nobelium (259) 2-8-18-32-32-8-2		103 Lr Lawrencium (266) 2-8-18-32-32-8-3																					

Atomic Number → 1

← Symbol H

Name → Hydrogen

← Atomic Weight 1.008

Electrons per shell → 1

State of matter (color of name)
GAS LIQUID SOLID UNKNOWN

Subcategory in the metal-metalloid-nonmetal trend (color of background)
Alkali metals Alkaline earth metals Transition metals Lanthanides Actinides Post-transition metals Metalloids Reactive nonmetals Noble gases Unknown chemical properties

Classification of elements

On the basis of physical properties, elements have been divided into the following three types:

- (i) Metals
- (ii) Non-metals
- (iii) Metalloids



With the exception of hydrogen, all elements that form positive ions by losing electrons during chemical reactions are called metals. Thus metals are electropositive elements with relatively low ionization energies. They are characterized by bright luster, hardness, ability to resonate sound and are excellent conductors of heat and electricity. Metals are solids under normal conditions except for Mercury.

Physical Properties of Metals

1. Metals are lustrous, malleable, ductile, good conductors of heat and electricity. Other properties include:
2. **State:** Metals are solids at room temperature with the exception of mercury, which is liquid at room temperature (Gallium is liquid on hot days).
3. **Luster:** Metals have the quality of reflecting light from their surface and can be polished e.g., gold, silver and copper.
4. **Malleability:** Metals have the ability to withstand hammering and can be made into thin sheets known as foils. For example, a sugar cube sized chunk of gold can be pounded into a thin sheet that will cover a football field.
5. **Ductility:** Metals can be drawn into wires. For example, 100 g of silver can be drawn into a thin wire about 200 meters long.
6. **Hardness:** All metals are hard except sodium and potassium, which are soft and can be cut with a knife.
7. **Valency:** Metals typically have 1 to 3 electrons in the outermost shell of their atoms.
8. **Conduction:** Metals are good conductors because they have free electrons. Silver and copper are the two best conductors of heat and electricity. Lead is the poorest conductor of heat. Bismuth, mercury and iron are also poor conductors
9. **Density:** Metals have high density and are very heavy. Iridium and osmium have the highest densities whereas lithium has the lowest density.
10. **Melting and Boiling Points:** Metals have high melting and boiling points. Tungsten has the highest melting and boiling points whereas mercury has the lowest. Sodium and potassium also have low melting points.

Non-metals

Elements that tend to gain electrons to form anions during chemical reactions are called non-metals. These are electronegative elements with high ionization energies. They are non-lustrous, brittle and poor conductors of heat and electricity (except graphite). Non-metals can be gases, liquids or solids.

Physical Properties of Nonmetals

- 1. Physical State:** Most of the non-metals exist in two of the three states of matter at room temperature: gases (oxygen) and solids (carbon). Only bromine exists as a liquid at room temperature.
- 2. Non-Malleable and Ductile:** Non-metals are very brittle, and cannot be rolled into wires or pounded into sheets.
- 3. Conduction:** They are poor conductors of heat(exception diamond) and electricity(exception graphite).
- 4. Luster:** These have no metallic luster and do not reflect light(exception iodine and graphite).
- 5. Melting and Boiling Points:** The melting points of non-metals are *generally* lower than metals, but are highly variable.
6. Seven non-metals exist under standard conditions as diatomic molecules: $\text{H}_2(\text{g})$, $\text{N}_2(\text{g})$, $\text{O}_2(\text{g})$, $\text{F}_2(\text{g})$, $\text{Cl}_2(\text{g})$, $\text{Br}_2(\text{l})$, $\text{I}_2(\text{s})$.
- 7. Non- metals have low densities**
 - Non- metals are generally light elements. That is why have low densities. For example, density of Sulphur is $2\text{g}/\text{cm}^3$.
- 8. Non- metals exists in different colors**

For example, Sulphur is yellow, phosphorus is white, yellow or red, graphite is black, chlorine is yellowish green, bromine is reddish brown while hydrogen, oxygen, nitrogen are colorless.

Metalloids

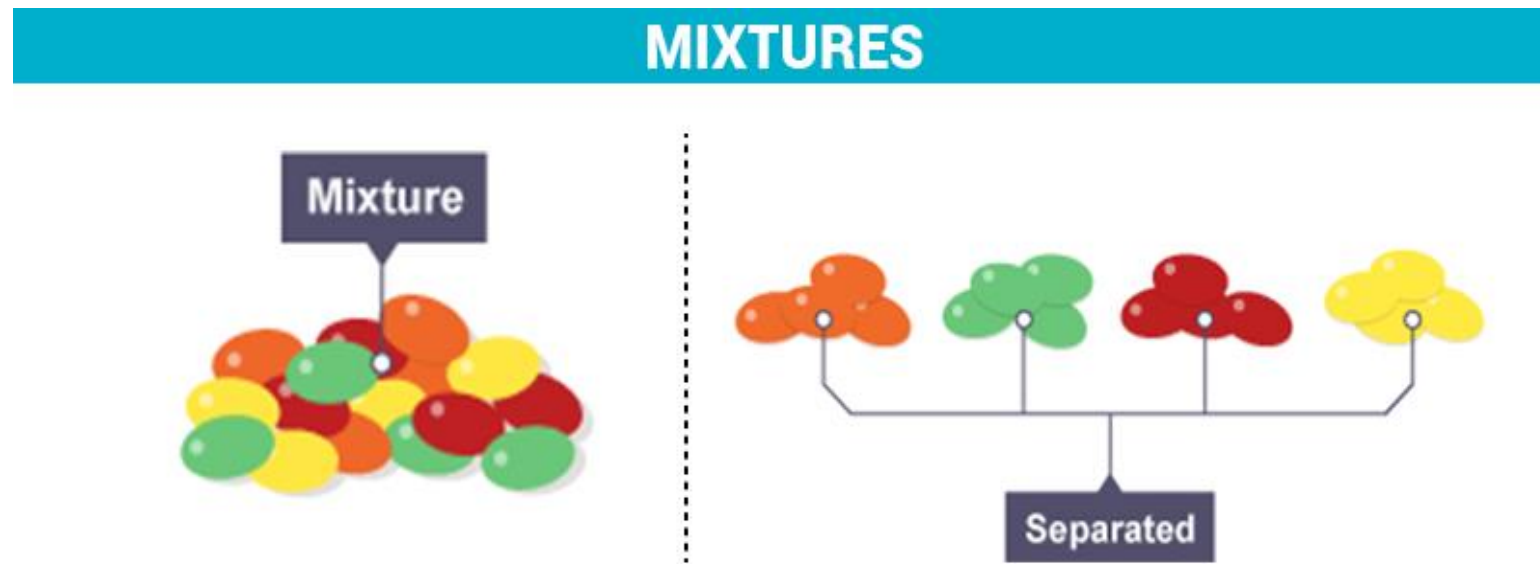
1. Metalloids have properties intermediate between the metals and nonmetals. Metalloids are useful in the semiconductor industry.
2. Metalloids are all solid at room temperature.
3. They can form alloys with other metals.
4. Some metalloids, such as silicon and germanium, can act as electrical conductors under the right conditions, thus they are called **semiconductors**.
5. Silicon for example appears lustrous, but is *not* malleable nor ductile (it is *brittle* - a characteristic of some nonmetals). It is a much poorer conductor of heat and electricity than the metals.
6. The physical properties of metalloids tend to be metallic, but their chemical properties tend to be non-metallic. The oxidation number of an element in this group can range from +5 to -2, depending on the group in which it is located.

Example of metalloids: Boron, Silicon, Germanium, Arsenic, Antimony, Tellurium, Astatine and Polonium.

Mixtures

In Chemistry when two or more substances mix with each other it results in the formation of a Mixture.

The result formed due to the combination of substances does not lose its individuality nor are they combined chemically. Mixtures are the one product of a mechanical blending or mixing of chemical substances such as elements and compounds.



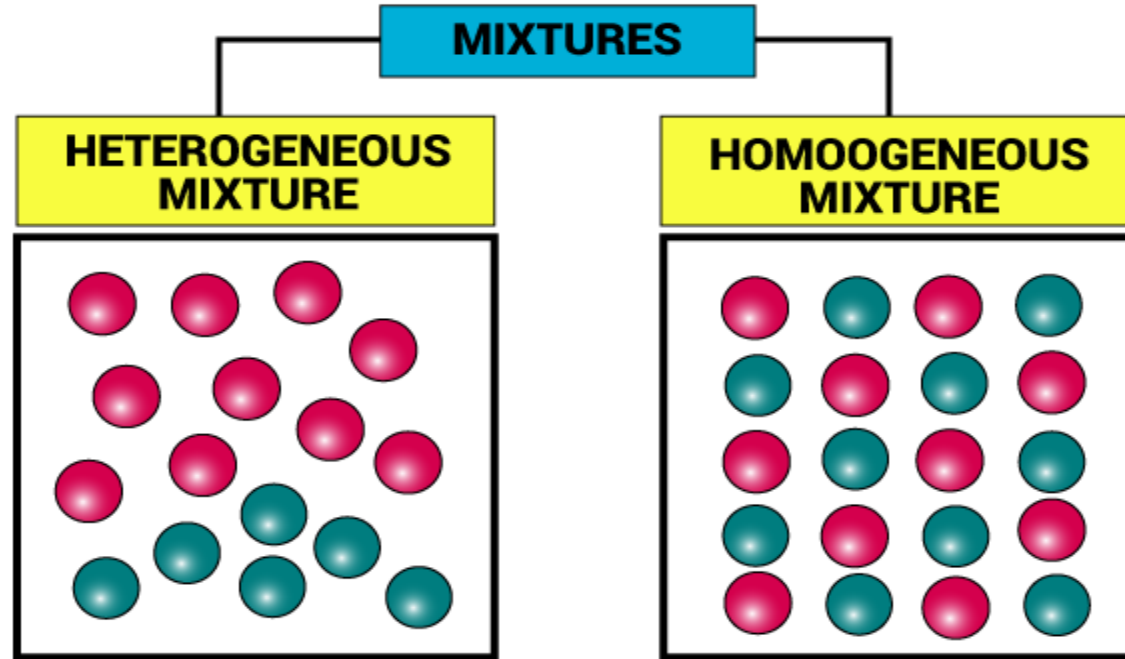
General Properties of Mixtures

1. Mixtures are made up of two or more substances that are not chemically combined with each other. The properties of mixtures are listed below.
2. The components of a mixture each keep their original properties.
3. The separation of components can be easily done.
4. The proportion of the components is variable.

Examples of Mixtures

- **Crude oil:** A mixture of organic compounds (mainly hydrocarbons)
- **Seawater:** A mixture of various salt and water.
- **Air:** a mixture of various gases like oxygen, carbon dioxide, nitrogen, argon, neon, etc.
- **Ink:** A mixture of coloured dyes.
- **Gunpowder:** A mixture of sulphur, potassium nitrate and carbon.

Types of Mixtures



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There are two main types of mixtures: [homogeneous mixtures and heterogeneous mixtures.](#)

1. What is a Heterogeneous Mixture?

A mixture of sand mixed with salt is an example of a heterogeneous mixture. Heterogeneous mixtures possess different properties and compositions in various parts i.e. the properties are not uniform throughout the mixture.

Examples of Heterogeneous mixtures – oil and water

2. What is a Homogeneous Mixture?

Sugar mixed with water is the most common example of a homogeneous mixture. [Homogeneous mixtures](#) can be defined as the mixtures which possess the same properties and combination throughout their mass.

Examples of Homogeneous mixtures – alloys, salt and water, alcohol in water, etc.

Characteristics of Mixtures

1. The constituents of a mixture are not present in a fixed ratio. The various characteristics of mixtures are discussed below.
2. There is no chemical force acting between the two or more substances that are mixed, but they still exist together.
3. They can either be heterogeneous or homogeneous in nature.
4. The proportions of the substances vary in an indefinite manner.
5. The properties of the mixture depending upon the individual components.
6. The constituents of the mixture can be separated by physical methods.
7. Boiling point and the melting point of the mixture depends upon the characteristic of the constituents.
8. During the formation of a mixture, there is no change in energy.
9. All the states of matter (solid, liquid, gases) can combine to form mixtures.

It can be concluded that almost everything in our vicinity is nothing but a mixture. **For example,** the food we eat is a mixture of ingredients, the atmospheric air we breathe is a combination of gases and the fuel we use in locomotives is a heterogeneous mixture.

Note:

Hydrogen chloride gas is a pure substance since it is chemical compound i.e. HCl(g) .

However, hydrochloric acid prepared by passing the gas into water is a homogeneous mixture. It is therefore, a solution.



Activity



Aim

To illustrate the concept of homogeneous and heterogeneous mixtures.

Materials required

Beakers, water, spatula, copper sulphate powder, potassium permanganate or common salt.

Method

- Let us divide the class into groups A, B, C and D.
- Group A takes a beaker containing 50 mL of water and one spatula full of copper sulphate powder. Group B takes 50 mL of water and two spatula full of copper sulphate powder in a beaker.
- Group C and D can take different amounts of copper sulphate and potassium permanganate or common salt(sodium chloride) and mix the given components to form a mixture.
- Report the observations on the uniformity in colour and texture.

Observation and discussion

- Both groups A and B have obtained homogeneous mixtures since the composition of these mixtures or solutions is uniform throughout.
- Although both the groups have obtained copper sulphate solutions but the intensity of colour of the two solutions is different. The intensity of blue colour in the solution obtained by group B which contains two spatula full of copper sulphate is much higher than the solution obtained by group A which contains one spatula full of copper sulphate.
- Both groups C and D have obtained heterogeneous mixtures since they not only have physically distinct boundaries but also their composition is not uniform.

Conclusion

- Soluble substances such as copper sulphate, common salt or sugar when dissolved in water form homogeneous mixtures, whose composition depend upon the amount of the substance dissolved.
- When two or more solids which do not react chemically are mixed, they always form heterogeneous mixtures.

In order to understand the difference between solution, colloid and suspension, let's perform the following activity.



Activity

Aim

To illustrate the difference between true solutions, suspensions and colloidal solutions.

Materials required

Beakers, glass rods, spatula, copper sulphate, chalk powder, wheat flour, milk, ink, torch.

Method

Let us again divide the class into four group – A, B, C and D.

- (a) Distribute the following samples to each group.
 - (i) Few crystals of copper sulphate to group A.
 - (ii) One spatula full of copper sulphate to group B.
 - (iii) Chalk powder or wheat flour to group C.
 - (iv) Few drops of milk or ink to group D.
- (b) Each group should add the given sample in water and stir properly using a glass rod. Are the particles in the mixture visible?
- (c) Direct a beam of light from a torch through the beaker containing the mixture and observe from the front. Was the path of the beam of light visible?
- (d) Leave the mixtures undisturbed for a few minutes (and set up the filtration apparatus in the meantime). Is the mixture stable or do the particles begin to settle after some time?
- (e) Filter the mixture. Is there any residue on the filter paper?

Observation and discussion

1. Group A and B

- (i) Group A and B both have got **true solutions**. However the intensity of the blue colour of solution obtained by group A which contains only a few crystals of copper sulphate is lower than that of the solution obtained by group B which contains one spatula full of copper sulphate. Thus, the intensity of colour depends upon the amount of substance dissolved.
 - (ii) In both these solutions, particles are not visible to the naked eye.
 - (iii) When the beam of light is passed through solution obtained by either group A or group B, the path of the beam of light was not visible. Thus, the particles of a true solution do not scatter light.
 - (iv) When the above solution is allowed to stand, the solution is stable and the particles of the solution do not settle down. Further, when the above solution is allowed to filter, the whole solution passes through the pores of the filter paper without leaving any residue on the filter paper.
- Thus, the particles of a true solution neither settle down on standing nor leave any residue when passed through a filter paper.

2. Group C

- (i) The group C has obtained a **suspension**.
- (ii) The particles of a freshly made suspension scatter light and hence are visible to the naked eye.
- (iii) When the above suspension is allowed to stand, it is not stable and the particles of the suspension settle down.
- (iv) When the above suspension is filtered, the particles of the chalk or the wheat flour being bigger than the pores of the filter paper remain as residue on the filter paper.

3. Group D

- (i) The group D has obtained a **colloidal solution**.
- (ii) When the above solution is allowed to stand, the particles of the colloidal solution like those of the true solution do not settle down.
- (iii) Since the particles of a colloidal solution are about 100 times bigger than those of the true solution, therefore, when a beam of light is passed through a colloidal solution, the particles of the colloidal solution scatter light and hence the path of the beam becomes visible.
- (iv) When the colloidal solution is passed through a filter paper, the particles of the colloidal solution like those of the true solution pass through the filter paper without leaving any residue on the filter paper.

Conclusion

- (i) The particles of a true solution are not visible even under a microscope. They do not settle down on standing and pass through the pores of the filter paper without leaving any residue.
- (ii) The particles of a suspension are visible to the naked eye. They settle down on standing and leave a residue on the filter paper.
- (iii) The particles of a colloidal solution scatter light and hence the path of light beam becomes visible when a beam of light is passed through it. A colloidal solution is stable and hence the particles of the colloidal solution do not immediately settle down on standing. They also pass through the pores of a filter paper.

Compounds

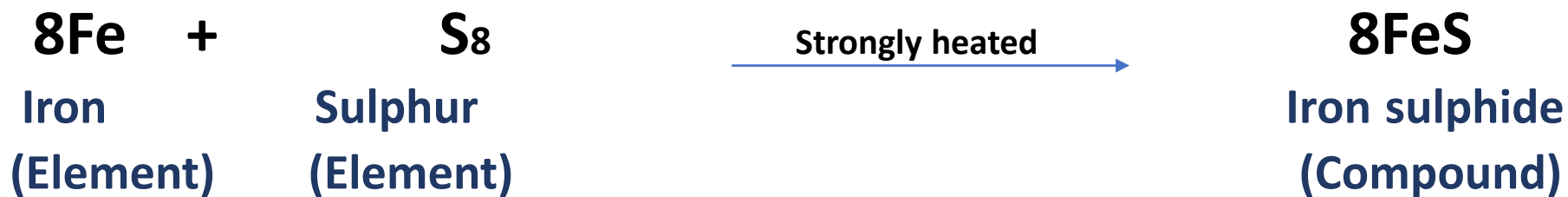
A compound is defined as a pure substance made up of two or more elements chemically combined together in a fixed proportion by mass.

Water(H_2O) is a compound made up of two elements, hydrogen(H) and oxygen(O), chemically combined together in a fixed proportion of 1:8 by mass.

Other examples of compounds are methane(CH_4), silica or sand (SiO_2), ammonia(NH_3), ammonium chloride(NH_4Cl), caustic soda(NaOH), washing soda(Na_2CO_3), lime or quick lime or calcium oxide(CaO), slaked lime or Calcium hydroxide [$\text{Ca}(\text{OH})_2$], potassium nitrate(KNO_3), copper sulphate(CuSO_4), potassium sulphate(K_2SO_4), potassium permanganate(KMnO_4), iron sulphide(FeS), hydrogen chloride(HCl), sulphuric acid(H_2SO_4), nitric acid(HNO_3), glucose($\text{C}_6\text{H}_{12}\text{O}_6$), sugar($\text{C}_{12}\text{H}_{22}\text{O}_{11}$), etc.

Properties of compounds

1. A compound is made up of the different elements combined together in a fixed proportion by mass.
2. A chemical compound is formed as a result of a chemical change(or reaction) and its properties are entirely different from those of its constituents. A chemical compound cannot be formed simply by mixing constituent elements. It is formed only when a chemical reaction takes place between them. For example, simply by mixing iron filings and sulphur powder, we get a mixture whose properties are the same as those of its constituents. However, when the mixture is heated strongly a chemical reaction occurs and a new compound called iron sulphide is formed.





Iron

+



Sulphur



Iron sulphide



The properties of iron sulphide thus formed are altogether different from those of its constituents, i.e. iron and Sulphur. For example, when dilute sulphuric acid is added to iron sulphide, a foul smelling gas (having the smell of rotten eggs) called hydrogen sulphide (H_2S) is evolved.

However, when iron reacts dilute sulphuric acid, hydrogen gas is formed which is a colourless, odourless combustible gas. This means that iron sulphide does not show the properties of iron present in it.

Similarly, water is not formed when hydrogen and oxygen are simply mixed together. It is formed only when an electric spark is passed through this mixture. The properties of water thus formed are altogether different from those of its constituent elements. For example, hydrogen is a combustible gas while oxygen is a supporter of combustion. Water on the other hand is neither combustible nor a supporter of combustion, it actually extinguishes the fire.

3. A compound cannot be separated into its constituents by simple physical means.

For example, water cannot be separated into its constituents i.e. hydrogen and oxygen elements by simple physical means such as filtration, evaporation, distillation, solvent extraction, or using a magnet. This separation can however, be done by chemical means i.e. by electrolysis (or passing electricity)

4. A compound has a definite molecular formula and fixed melting point and boiling point.
5. Energy in form of heat or light is usually evolved or absorbed when a compound is formed.

For example, formation of carbon dioxide(CO_2) by burning of coke in air or oxygen is always accompanied by evolution of heat.

6. A compound is a homogeneous substance.

For example, if we take water in a glass or a beaker, every drop of it looks alike i.e. it is homogeneous and will show the same properties.

Important notes

1. Compounds are homogeneous but mixture may be homogeneous or heterogeneous.

Homogeneity alone cannot be used to decide between a mixture and a compound. For example, solutions of common salt, sugar, copper sulphate, etc. in water are homogeneous. But still they are regarded as mixtures because of the following reasons:

- (i) The common salt solution shows the properties of its constituents i.e. common salt and water.
- (ii) The constituents of the salt solution i.e. common salt and water can be separated by the physical process of distillation.
- (iii) The composition of the salt solution is variable. In other words, different amounts of salt can be dissolved in the same amount of water to get solutions having different compositions. Thus, a salt solution does not have a definite formula.
- (iv) Salt solution does not have a fixed boiling point.

2. Non-stoichiometric compounds

Molecules (constituting units of compounds) which have definite atomic composition are referred as stoichiometric compounds. For example, **CH₄**, **H₂O**, **NH₃**, etc.

Certain materials in which atomic composition is variable are called non-stoichiometric compounds. For example, the composition of cuprous sulphide may vary from **Cu_{1.7}S** to **Cu₂S**.

Non- stoichiometry is common feature of sulphides and oxides of transition metals.

Types of compounds

All the compounds may be divided into following two categories:

1. **Organic compounds**
2. **Inorganic compounds**

Organic compounds are the compounds containing carbon and few other elements like hydrogen, oxygen, nitrogen, sulphur, halogen, etc. These were originally obtained only from animals and plants.

Inorganic compounds are the compounds containing any two or more elements out of more than 118 elements known so far. These are usually obtained from minerals and rocks.

Criteria of purity of compounds

A solid compound is said to be pure if it has a sharp melting point i.e. whole of the solid melts within a range of **0.5 °C** or **0.5 K**. Impure solids on the other hand have lower melting points than the pure solids because impurities lower or depress the melting points.

A liquid compound is said to be pure if it has a fixed boiling point i.e. whole the liquid distills at a fixed temperature.

Impure liquids on the other hand have higher boiling points than the pure liquids. This is because impurities always raise the boiling points. For example, pure water boils at **100 °C** (373K) at one atmosphere pressure. However, sea water is impure because it contains a number of dissolved salts. Thus, the boiling point of sea water is always higher than **100 °C** (373 K).

Question

List the following ‘ waters in order of increasing purity:

Ocean water, rainwater and drinking water

Explanation:

Ocean water is the least pure because it contains a large amount of dissolved compounds. That is why it is not drinkable and cannot be used for crop irrigation. Drinking water also contains chlorine and some dissolved compounds but not nearly as much as ocean water. Rain water is the purest of three but still contains some dissolved gases from the air.

Solution

A solution is a homogeneous mixture of two or more components in which the particle size is smaller than 1 nm.

Common examples of solutions are the sugar in water and salt in water solutions, soda water, etc. In a solution, all the components appear as a single phase. There is particle homogeneity i.e. particles are evenly distributed. This is why a whole bottle of soft drink has the same taste throughout.

It may be noted that all mixtures are not solutions. If a mixture is to be called a solution. It must satisfy the following two conditions.

1. Components should be non reacting.
2. Should be homogeneous.

Solution Examples

- Air is a homogeneous mixture of gases. Here both the solvent and the solute are gases.
- Sugar syrup is a solution where sugar is dissolved in water using heat. Here, water is the solvent and sugar is the solute.
- Tincture of iodine, a mixture of iodine in alcohol. Iodine is the solute whereas alcohol is the solvent.

ALLOYS

Alloys are metal mixtures with other elements, the precise combination of which is governed by the properties required. In general, alloys are considered metallic in nature, i.e. they have good thermal and electrical conductivity. **An alloy is a homogeneous mixture of two or more metals or a metal and a nonmetal.**

Alloys have a wide variety of uses. The creation of alloys actually reduces the material's price and maintains the metal's desired properties. The combination of metals improves material properties and offers resistance to oxidation and mechanical strength.

The **brass alloy(30% zinc and 70% copper)** is used for many home decorative items and consists of copper and zinc. The combination of these metals makes the brass solid and robust. The door buttons and high-quality locks all make of brass.



Components of solution

Solutions have two components, one is solvent and the other is solute and such solution is called a binary solution.

The component that dissolves the other component is called the solvent.

The component(s) that is/are dissolved in the solvent is/are called solute(s).

Generally solvent is present in major proportion compared to the solute. The amount of solute is lesser than the solvent. The solute and solvent can be in any state of matter i.e. solid, liquid and gas.

Solutions that are in the liquid state consist of a solid, liquid or gas dissolved in a liquid solvent. Alloys and air are examples of solid and gaseous solutions respectively.

(a) Solvent: The component of a solution which dissolves the other component in itself, is called solvent. For example, a solution of sugar in water is a solid in liquid solution. In this solution, sugar is the solute and water is the solvent.

(b) Solute: The component of a solution which dissolves in the solvent, is called solute.

For example, in solution of iodine in alcohol, known as 'tincture of iodine', iodine is the solute.

Different properties of solutions are as follows:

1. It is a homogeneous mixture.
2. Its particles are too tiny and have a diameter less than 1 nm.
3. The particles are not visible to naked eyes.
4. Particles don't scatter a beam of light passing through it and hence the path of the light is not visible.
5. Solutes are inseparable from the mixture and do not sediment. A solution is stable.
6. The components of a mixture cannot be separated using filtration.

A solution in which water acts as the solvent is called an **aqueous solution**.

The common examples of aqueous solutions are

- (i) Common salt dissolved in water.
- (ii) Sugar dissolved in water.
- (iii) Acetic acid dissolved in water(called vinegar)

Most of the substances are soluble in water. That is why water is sometimes called a 'Universal solvent'. However, all substances do not dissolve in water.

If a solute is dissolved in any solvent such as benzene, alcohol, ether, carbon disulphide, carbon tetrachloride, acetone, benzene etc., the solution obtained is called a **non aqueous solution**.

The common examples of non aqueous solutions are

- (i) Bromine dissolved in chloroform.
- (ii) Iodine dissolved in carbon tetrachloride.
- (iii) Sulphur dissolved in carbon disulphide.

Types of Solution

Liquid solutions, such as sugar in water, are the most common kind, but there are also solutions that are gases or solids. Any state of matter (solid, liquid, or gas) can act both as a solute or as a solvent during the formation of a solution. Therefore depending upon the physical states of solute and solvent, we can classify in nine different types of solutions.

S.No	Types of Solution	Solute	Solvent	Examples
1	Solid-solid	solid	solid	Alloys like brass, bronze etc.
2	Solid-liquid	solid	liquid	The solution of sugar, salt etc in water.
3	Solid-gas	solid	gas	Sublimation of substances like iodine, camphor etc into the air.
4	Liquid-solid	liquid	solid	Hydrated salts, mercury in amalgamated zinc etc.
5	Liquid-liquid	liquid	liquid	Alcohol in water, benzene in toluene
6	Liquid-gas	liquid	gas	Aerosol, water vapour in the air.
7	Gas-solid	gas	solid	Hydrogen absorbed in palladium
8	Gas-liquid	gas	liquid	Aerated drinks
9	Gas-gas	gas	gas	A mixture of gases etc

CONCENTRATION TERMS

Concentration is a very common concept used in chemistry and related fields. It is the measure of how much of a given substance there is mixed with another substance. This can apply to any sort of chemical mixture, but most frequently is used in relation to solutions, where it refers to the amount of solute dissolved in a solvent. **To concentrate a solution, one must add more solute, or reduce the amount of solvent (for instance, by selective evaporation).** By contrast, to dilute a solution, one must add more solvent, or reduce the amount of solute. There exists a concentration at which no further solute will dissolve in a solution. At this point, the solution is said to be saturated. If additional solute is added to a saturated solution, it will not dissolve. Instead, phase separation will occur, leading to either coexisting phases or a suspension. The point of saturation depends on many variables such as ambient temperature and the precise chemical nature of the solvent and solute. **Concentration may be expressed both qualitatively ('informally') and quantitatively ('numerically').**

1. Qualitative notation: Qualitatively, solutions of relatively low concentration are described using adjectives such as **“dilute,”** or **“weak,”** while solutions of relatively high concentration are described as **“concentrated,”** or **“strong.”** As a rule, the more concentrated a chromatic solution is, the more intensely coloured it is.

These glasses containing red dye demonstrate qualitative changes in concentration. The solutions on the left are **“weaker”** (or more dilute), compared to the **“stronger”** (or more concentrated) solutions on the right.

2. Quantitative notation: Quantitative notation of concentration is far more informative and useful from a scientific point of view. There are a number of different ways to quantitatively express concentration; the most common are listed below.

Note: Many units of concentration require measurement of a substance’s volume, which is variable depending on ambient temperature and pressure. Unless otherwise stated, all the following measurements are assumed to be at standard state temperature and pressure (that is, 25 degrees Celsius at 1 atmosphere).

Concentration of a Solution

The concentration of a solution is the amount of solute present in a given quantity of the solution.

The solution having small amount of solute is called **Dilute Solution**.

The solution having large amount of solute is called **Concentrated Solution**.

The concentration of a solution is defined as mass of solute in grams present in 100 g of the solution.

For example: 10% solution of NaCl (10 g of NaCl + 90 g of water)

The amount of solute dissolved in a given mass or volume of solvent.

$$\text{Concentration of solution} = \frac{\text{Amount of solute}}{\text{Amount of solution}}$$

Methods of Expressing the Concentration of a Solution

The concentration of a solution can be expressed in a number of ways. The important methods are:

1. Mass percent
2. Volume percent
3. Mass/Volume percent
4. Molarity
5. Normality
6. Formality
7. Molality
8. Mole fraction

Percentage by Mass or Mass Percentage (w/w):

This method is used for a solid in a liquid solution. The mass of solute in gram dissolved in the solvent to form 100 grams of the solution is called percentage by mass. The ratio of the mass of solute to the mass of the solution is called a mass fraction.

For example, if a solution is described by 10% glucose in water by mass, it means that 10 g of glucose is dissolved in 90 g of water resulting in a 100 g solution.

Concentration described by mass percentage is commonly used in industrial chemical applications. For example, a commercial bleaching solution contains a 3.62 mass percentage of sodium hypochlorite in water.

$$\text{Percentage by mass} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$
$$\text{Percentage by mass} = \frac{\text{Mass of Solute}}{\text{Mass of solute} + \text{Mass of solvent}} \times 100$$

Question:

A solution contains 40 g of common salt in 320 g of water. Calculate the concentration in terms of mass by mass percentage of the solution

Solution

Mass of solute (salt) = 40 g

Mass of solvent (water) = 320 g

Also, Mass of solution = Mass of solute + Mass of solvent

$$= (40 \text{ g} + 320 \text{ g})$$

$$= 360 \text{ g}$$

Mass percentage of solution = $\left[\frac{\text{Mass of solute}}{\text{Mass of solution}} \right] \times 100$

$$= (40/360) \times 100 = 11.1\%$$

Question:

What is the weight percentage of urea solution in which 10 gm of urea is dissolved in 90 gm water.

Solution

Mass of solute (urea) = 10 gm

Mass of solvent (water) = 90 gm

Also, Mass of solution = Mass of solute + Mass of solvent

$$= (10 \text{ gm} + 90 \text{ gm})$$

$$= 100 \text{ gm}$$

Mass percentage of solution = $\left[\frac{\text{Mass of solute}}{\text{Mass of solution}} \right] \times 100$

$$= (10/100) \times 100 = 10\% \text{ urea solution}$$

Percentage by Volume (V/V):

This method is used for liquid in a liquid solution. For example, a 10% ethanol solution in water means that 10 mL of ethanol is dissolved in 90 mL water such that the total volume of the solution is 100 mL.

$$\text{Percentage by volume} = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$$

$$\text{Percentage by volume} = \frac{\text{Volume of Solute}}{\text{Volume of solute} + \text{Volume of solvent}} \times 100$$

Percentage by Mass by Volume (w/V):

It is the mass of solute dissolved in 100 mL of the solution. This method is commonly used in medicine and pharmacy.

$$\text{Percentage mass by volume} = \frac{\text{Mass of solute}}{\text{Volume of solution}} \times 100$$

Question:

In a solution, there is 111.0 mL (110.605 g) solvent and 5.24 mL (6.0508 g) solute present in a solution. Find the mass percent, volume percent and mass/volume percent of the solute.

Solution

Mass Percent

$$= (\text{Mass of Solute}) / (\text{Mass of Solution}) \times 100\%$$

$$= (6.0508\text{g}) / (110.605\text{g} + 6.0508\text{g}) \times 100\%$$

$$= (0.0518688312) \times 100\%$$

$$= 5.186883121\%$$

$$\text{Mass Percent} = 5.186\%$$

Volume Percent

$$= (\text{Volume of Solute}) / (\text{Volume of Solution}) \times 100\%$$

$$= (5.24\text{mL}) / (111.0\text{mL} + 5.24\text{mL}) \times 100\%$$

$$= (0.0450791466) \times 100\%$$

$$= 4.507914659\%$$

$$\text{Volume Percent} = 4.51\%$$

Mass/Volume Percent

$$= (\text{Mass of Solute}) / (\text{Volume of Solution}) \times 100\%$$

$$= (6.0508\text{g}) / (111.0\text{mL} + 5.24\text{mL}) \times 100\%$$

$$= (0.0520) \times 100\%$$

$$= 5.205\%$$

$$\text{Mass/Volume Percent} = 5.2054\%$$

Practice questions:

1. If 110g of salt is present in 550 g of solution, calculate the concentration.
2. What is the concentration of a solution which contains 16 g of urea in 120 g of solutions?
3. How much water should be added to 15 g of salt to obtain 15% salt solution?
4. A solution contains 5.6 ml of alcohol mixed with 76 ml of water. Calculate the concentration of the solution.
5. How much water should be added to 12 ml of alcohol to obtain 12% alcohol solution?
6. If 25 ml of acetone is present in 150 ml of its aqueous solution, calculate the concentration.

Answers

1. 20%
2. 13.33%
3. 100 ml
4. 6.8627%
5. 88 ml
6. 16.66%

Molarity (M) : Molarity of a solution is the *number of moles of the solute per litre of solution (or number of millimoles per ml. of solution)*. Unit of molarity is ***mol/litre*** or ***mol/dm³*** For example, a molar (1M) solution of sugar means a solution containing 1 mole of sugar (*i.e.*, 342 g or 6.02×10^{23}

molecules of it) per *litre* of the solution. Solutions in term of molarity generally expressed as,

1M = Molar solution, **2M** = Molarity is two,

$\frac{M}{2}$ or 0.5 M = Semimolar solution,

$\frac{M}{10}$ or 0.1 M = Decimolar solution,

$\frac{M}{100}$ or 0.01 M = Centimolar solution

$\frac{M}{1000}$ or 0.001 M = Millimolar solution

- Molarity is most common way of representing the concentration of solution.
- Molarity is depend on temperature as, $M \propto \frac{1}{T}$
- When a solution is diluted (x times), its molarity also decreases (by x times)

Mathematically molarity can be calculated by following formulas,

$$(i) M = \frac{\text{No. of moles of solute (n)}}{\text{Vol. of solution in litres}},$$

$$(ii) M = \frac{\text{Wt. of solute (in gm) per litre of solution}}{\text{Mol. wt. of solute}}$$

$$(iii) M = \frac{\text{Wt. of solute (in gm)}}{\text{Mol. wt. of solute}} \times \frac{1000}{\text{Vol. of solution in ml.}}$$

$$(iv) M = \frac{\text{No. of millimoles of solute}}{\text{Vol. of solution in ml}}$$

$$(v) M = \frac{\text{Percent of solute} \times 10}{\text{Mol. wt. of solute}}$$

$$(vi) M = \frac{\text{Strength in } g l^{-1} \text{ of solution}}{\text{Mol. wt. of solute}}$$

$$(vii) M = \frac{10 \times \text{Sp. gr. of the solution} \times \text{Wt.}}{\text{Mol. wt. of the solute}}$$

(viii) If molarity and volume of solution are changed from M_1, V_1 to M_2, V_2 . Then,

$$M_1 V_1 = M_2 V_2 \text{ (Molarity equation)}$$

Molarity of dilution:

Before dilution

After dilution

$$M_1V_1 = M_2V_2$$

Molarity of mixing:

Let there be three samples of solution (containing same solvent and solute) with their molarity M_1 , M_2 , M_3 and volumes V_1 , V_2 , V_3 respectively. These solutions are mixed; molarity of mixed solution may be given as:

$$M_1V_1 + M_2V_2 + M_3V_3 = M_R(V_1 + V_2 + V_3)$$

Where M_R = resultant molarity

$V_1 + V_2 + V_3$ = resultant volume

The equation for calculating molarity is the ratio of the moles of solute whose molarity is to be calculated and the volume of solvent used to dissolve the given solute.

$$M = \frac{n}{V}$$

Here,

M is the molality of the solution that is to be calculated

n is the number of moles of the solute

V is the volume of solution given in terms of liters

Question:

A solution is prepared by bubbling 1.56 grams of hydrochloric acid in water. Here, the volume of the solution is 26.8 ml. Calculate the molarity of the solution.

Solution

The chemical formula of hydrochloric acid = HCl

The chemical formula for Water = H₂O

The molecular weight of HCl = $35.5 \times 1 + 1 \times 1 = 36.5$ moles/gram

The molecular weight of H₂O = $1 \times 2 + 16 \times 1 = 18$ moles/gram

Given, mass of hydrochloric acid in the solution = 1.56 g

The number of moles of hydrochloric acid =

$$n_T = \frac{\text{mass in grams}}{\text{molecular weight}}$$

Now, given volume of the solution = 26.8 mL

$$n_T = \frac{1.56}{36.5} = 4.27 \times 10^{-2} \text{ mole}$$

Expressing the volume in terms of liters,

$$\text{volume} = \frac{26.8}{1000} = 2.68 \times 10^{-2} \text{ Litre}$$

Now, we calculate the molarity of the solution using the formula given above.

$$\text{Molarity} = \frac{\text{Number of moles of element}}{\text{volume of solution in litres}}$$

The molarity of the solution is 1.59 M.

$$= \frac{4.27 \times 10^{-2}}{2.68 \times 10^{-2}} = 1.59 M$$

Question:

A 1.5L solution is composed of 0.25g NaCl dissolved in water. Find its molarity.

Solution

$$\text{Moles of NaCl} = (0.25\text{g}) / (22.99\text{g} + 35.45\text{g}) = 0.004277 \text{ mol NaCl}$$

Molarity

$$= (\text{Moles of Solute}) / (\text{Liters of Solution})$$

$$= (0.004277 \text{ mol NaCl}) / (1.5\text{L})$$

$$= 0.002851 \text{ M}$$

$$\text{Molarity} = 0.0029\text{M}$$

Question:

Determine the molarity for each of the following solutions:

- a. 0.444 mol of CoCl_2 in 0.654 L of solution
- b. 98.0 g of phosphoric acid, H_3PO_4 , in 1.00 L of solution
- c. 0.2074 g of calcium hydroxide, $\text{Ca}(\text{OH})_2$, in 40.00 mL of solution
- d. 10.5 kg of $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ in 18.60 L of solution
- e. 7.0×10^{-3} mol of I_2 in 100.0 mL of solution
- f. 1.8×10^4 mg of HCl in 0.075 L of solution

Answer a

0.679 *M*

Answer b

1.00 *M*

Answer c

0.06998 *M*

Answer d

1.75 *M*

Answer e

0.070 *M*

Answer f

6.6 *M*

Questions

1. There is about 1.0 g of calcium, as Ca^{2+} , in 1.0 L of milk. What is the molarity of Ca^{2+} in milk?

Answer: 0.025 M

2. What volume of a 1.00-*M* $\text{Fe}(\text{NO}_3)_3$ solution can be diluted to prepare 1.00 L of a solution with a concentration of 0.250 *M*?

Answer: 0.250 L

3. If 0.1718 L of a 0.3556-*M* $\text{C}_3\text{H}_7\text{OH}$ solution is diluted to a concentration of 0.1222 *M*, what is the volume of the resulting solution?

Answer: 0.5000 L

$$\text{Normality (N)} = \frac{\text{Number of gram equivalents of solute}}{\text{Volume of Solution in L}}$$

The relation between normality and molarity.

- $N \times \text{Eq.Wt} = \text{Molarity} \times \text{Molar mass}$
- $N = \text{Molarity} \times \text{Valency}$
- $N = \text{Molarity} \times \text{Number of H}^+ \text{ or OH}^- \text{ ion.}$

$$N = \frac{x \times d \times 10}{\text{Equivalent Mass of Solute}}$$

Number of gram equivalent of solute = Mass of solute in gram/ equivalent weight of solute

Equivalent weight of solute (E) = Molar mass of solute/ Valence factor

Valence factor for base = acidity of base

Valence factor for acid = basicity of acid

Valence factor for element = valency

Thus, if one gram equivalent of a solute is present in one litre of the solution, the concentration of the solution is said to be *one normal*.

1N = Normal = One gram equivalent of the solute per litre of solution = Normality is 1

N/2 = Seminormal = 0.5 g equivalent of the solute per litre of solution = Normality is 0.5

N/10 = Decinormal = 0.1 g equivalent of the solute per litre of solution = Normality is 0.1

N/100 = Centinormal = 0.01 g equivalent of the solute per litre of solution = Normality is 0.01

N/1000 = Millinormal = 0.001 g equivalent of the solute per litre of solution = Normality is 0.001

Relationship between normality and molarity:

We know that

Molarity $\times$ Molecular mass = Strength of solution (g/L)

Similarly,

Normality $\times$ Equivalent mass = Normality of the solution (g/L)

Hence,

Molarity $\times$ Molecular mass = Normality $\times$ Equivalent mass

So, Normality = $n \times$ Molarity

Let 'd' is the density of solution in g/mL

and

x is the percentage of the solute by mass.

Then,

$$N = \frac{x \times d \times 10}{\text{Equivalent Mass of Solute}}$$

Question:

In the following reaction calculate and find the normality when it is 1.0 M H_3PO_4



Solution

If we look at the given reaction we can identify that only two of the H^+ ions of H_3AsO_4 react with NaOH to form the product. Therefore, the two ions are 2 equivalents. In order to find the normality, we will apply the given formula.

$$N = \text{Molarity (M)} \times \text{number of equivalents}$$

$$N = 1.0 \times 2 \text{ (replacing the values)}$$

Therefore, normality of the solution = 2.0.

Question:

Calculate the normality of 0.321 g sodium carbonate when it is mixed in a 250 mL solution.

Solution

First, you have to know or write down the formula for sodium carbonate. Once you do this you can identify that there are two sodium ions for each carbonate ion. Now solving the problem will be easy.

N of 0.321 g sodium carbonate

$$N = \text{Na}_2\text{CO}_3 \times (0.321 \text{ g} / 105.99 \text{ g}) \times (2 \text{ eq}/1 \text{ mol})$$

$$N = 0.00605 \text{ eq}/0.2500 \text{ L}$$

$$\mathbf{N = 0.02422 \text{ N}}$$

Question:

What is the normality of the following?

a. 0.1381 M NaOH

b. 0.0521 M H₃PO₄

Solution

a. $N = 0.1381 \text{ mol/L} \times (1 \text{ eq/1mol}) = 0.1381 \text{ eq/L} = 0.1381 \text{ N}$

b. $N = 0.0521 \text{ mol/L} \times (3 \text{ eq/1mol}) = 0.156 \text{ eq/L} = 0.156 \text{ N}$

Formality

It is the number of formula mass in grams present per litre of solution. In case formula mass is equal to molecular mass, formality is equal to molarity. Like molarity and normality, the formality is also dependent on temperature. It is used for ionic compounds in which there is no existence of molecule. Mole of ionic compounds is called formole and molarity as formality.

$$\text{Formality} = \frac{\text{Weight of solute (gm)}}{\text{Formula weight of solute}} \times \frac{1}{\text{Volume of solutions (L)}}$$

$$F = \frac{w}{f} \times \frac{1}{V (L)} \cdot (i)$$

Where,

w = weight of solute,

f = formula weight of solute

V = volume of solution

n_f = no. of gram formula weight

Question:

CH_3COOH exists as dimer in benzene. 1.2 gm of the acid was dissolved and the volume was made up to one litre using benzene. What is the formality?

Solution

Molar mass of $\text{CH}_3\text{COOH} = 60$

Formula weight of the associated molecule of the acid $= 2 \times 60 = 120$

Weight of $\text{CH}_3\text{COOH} = 1.2 \text{ gm}$

Volume of solution $= 1 \text{ L}$

Formality $= 1.2/120 \times 1/1 = 0.01 \text{ F}$

Molality

Molality is a solution property and is defined as the number of solvent moles per kilogram. Molality's SI unit is mol/kg. A solution with a 3 molar/kg molality is often defined as "3 molal" or "3 m." However, it is now preferred following the unit SI system, mol/kg or a similar SI unit.

Molality Formula

$$\text{Molality } (M) = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kgs}}$$

$$\text{Molality } (M) = \frac{g \times 1000}{W \times m}$$

Relation Between Molarity And Molality:

Let the mass of given solute be W.

Let the volume of the solution be V.

Let the molality be m.

Let the molar mass of solute be M'.

Let the Molarity be M.

Let the weight of the solvent be W'.

$$\text{Therefore the Molarity, } M = \frac{W}{M'} * \frac{1000}{V} \dots\dots\dots(1)$$

$$\text{Therefore the molality, } m = \frac{W}{M'} * \frac{1000}{W'} \dots\dots\dots(2)$$

$$\text{Density, } d = \frac{M}{V}$$

$$= \frac{W+W'}{V}$$

From (1)

$$V = \frac{W*1000}{MM'} \dots\dots\dots (3)$$

From (2)

$$W' = \frac{W*1000}{mM'}$$

Therefore,

$$W+W' = W + \frac{W*1000}{mM'} \frac{1000}{M'} \left[\frac{mM'}{1000m} + \frac{1}{m} \right] W' \dots\dots\dots(4)$$

By dividing equation (4) by (3) we get,

$$\frac{d}{M} = \frac{1}{m} + \frac{M'}{1000} \frac{1}{m} = \frac{d}{M} - \frac{M'}{1000} \quad m = \frac{M*1000}{(d*1000)-M(M')}$$

Hence the relation between molality and Molarity proved.

Question:

Calculate the molality of a solution where 5 grams of toluene (C_7H_8) is dissolved in 225 grams of benzene (C_6H_6).

Solution

Toluene: Molecular weight = $C_7H_8 = 7 \times 12 + 1 \times 8 = 92 \text{ grams/mole}$

Using the formula: $\text{Moles of Toluene} = \frac{\text{Mass in grams}}{\text{Molecular weight}} = \frac{5.0 \text{ grams}}{92 \text{ grams/mole}} = 0.054 \text{ mole.}$

So, the mole of toluene is 0.054 mole.

Now calculate the kilogram of solvent. $\frac{225 \text{ grams of Benzene}}{1000} = 0.225 \text{ kilogram}$

As the final step, calculate the molality using the formula.

$$\text{Molality (m)} = \frac{\text{Moles of Toluene}}{\text{Mass of Benzene in grams}}$$

$$\text{Molality (m)} = \frac{0.054 \text{ moles}}{0.225 \text{ kg}} \quad \text{Molality (m)} = 0.24 \text{ m}$$

Mole Fraction

If the solution has a solvent and the solute, a mole fraction gives a concentration as the ratio of moles of one component to the total moles present in the solution. It is denoted by x . Suppose we have a solution containing A as a solute and B as the solvent. Let n_A and n_B be the number of moles of A and B present in the solution respectively. So mole fractions of A and B are given as:

$$x_A = \frac{n_A}{n_A + n_B}$$
$$x_B = \frac{n_B}{n_A + n_B}$$

Calculations with Mole fraction and Mole Percent

Mole Fraction in Mixtures

A mixture of gases was formed by combining 6.3 moles of O_2 and 5.6 moles of N_2 . What is the mole fraction of nitrogen in the mixture?

First, we must find the total number of moles with $n_{\text{total}} = n_{N_2} + n_{O_2}$.

$$n_{\text{total}} = 6.3 \text{ moles} + 5.6 \text{ moles} = 11.9 \text{ moles.}$$

Next we must divide the moles of N_2 by the total number of moles:

$$x(\text{mole fraction}) = \left(\frac{\text{moles } N_2}{\text{moles } N_2 + \text{moles } O_2} \right) = \left(\frac{5.6 \text{ moles}}{11.9 \text{ moles}} \right) = 0.47$$

The mole fraction of nitrogen in the mixture is 0.47.

Mole Fraction in Solutions

Mole fraction can also be applied in the case of solutions. For example, 0.100 moles of NaCl are dissolved in 100.0 mL of water. What is the mole fraction of NaCl?

We are given the number of moles of NaCl, but the volume of water. First, we convert this volume to a mass by using the density of water (1.00 g/mL), and then we convert this mass to moles of water:

$$100 \text{ mL } H_2O \times \left(\frac{1.0g}{1mL} \right) = 100.0 \text{ g } H_2O \times \left(\frac{1 \text{ moles}}{18.0g} \right) = 5.55 \text{ moles } H_2O$$

With this information we can find the total number of moles present: $5.55 + 0.100 = 5.65$ moles. If we divide the moles of NaCl by the total number of moles, we find the mole fraction of this component:

$$x = \left(\frac{0.100\text{moles}}{5.65\text{moles}} \right) = 0.0176$$

Mole Fraction with Multi-Component Mixtures

Mole fractions can also be found for mixtures that are formed from multiple components. These are treated no differently than before; again, the total mole fraction of the mixture must be equal to 1.

For example, a solution is formed by mixing 10.0 g of pentane (C_5H_{12}), 10.0 g of hexane (C_6H_{14}) and 10.0 g of benzene (C_6H_6). What is the mole fraction of hexane in this mixture?

We must first find the number of moles present in 10.0 g of each component, given their chemical formulas and molecular weights. The number of moles for each is found by dividing its mass by its respective molecular weight. We find that there are 0.138 moles of pentane, 0.116 moles of hexane, and 0.128 moles of benzene.

We can find the total number of moles by taking the sum of all the moles: $0.138 + 0.116 + 0.128 = 0.382$ total moles. If we divide moles of hexane by the total moles, we calculate the mole fraction:

$$x = \left(\frac{0.116 \text{ moles}}{0.382 \text{ moles}} \right) = 0.303$$

Mole Fraction from Molality

Mole fraction can also be calculated from molality. If we have a 1.62 m solution of table sugar ($\text{C}_6\text{H}_{12}\text{O}_6$) in water, what is the mole fraction of the table sugar?

Since we are given molality, we can convert it to the equivalent mole fraction, which is already a mass ratio; remember that molality = moles solute/kg solvent. Given the definition of molality, we know that we have a solution with 1.62 moles of sugar and 1.00 kg (1000 g) of water. Since we know the number of moles of sugar, we need to find the moles of water using its molecular weight:

$$1000 \text{ g H}_2\text{O} \times \left(\frac{1 \text{ mole}}{18.0 \text{ g}} \right) = 55.5 \text{ moles H}_2\text{O}$$

The total number of moles is the sum of the moles of water and sugar, or 57.1 moles total of solution. We can now find the mole fraction of the sugar:

$$x = \left(\frac{1.62 \text{ moles sugar}}{57.1 \text{ moles solution}} \right) = 0.0284$$

Mole Fraction from Mass Percent

The mole fraction can also be calculated from a mass percent. What is the mole fraction of cinnamic acid that has a mass percent of 50.00% urea in cinnamic acid? The molecular weight of urea is 60.16 g/mol and the molecular weight of cinnamic acid is 148.16 g/mol.

First, we assume a total mass of 100.0 g, although any mass could be assumed. This means that we have 50.0 g of urea and 50.0 g of cinnamic acid. We can then calculate the moles present by dividing each by its molecular weight. We have 0.833 moles urea and 0.388 moles cinnamic acid, so we have 1.22 moles total.

To find the mole fraction, we divide the moles of cinnamic acid by total number of moles:

$$x = \left(\frac{.388 \text{ moles cinnamic acid}}{1.22 \text{ moles solution}} \right) = 0.318$$

The mole fraction for cinnamic acid is 0.318.

Demal unit (D) : The concentrations are also expressed in “Demal unit”. One demal unit represents one mole of solute present in one litre of solution at 0°C